Determine the pH of a solution made by dissolving 5.0 g (2 s.f.) of ~~acid or base?~~ acidic or base? $\boxed{\text{strontium hydroxide}}$ in 100 mL (1 s.f.) of water.

① acidic / basic → pH
→ pOH
② strong or weak
↳ no equilibrium
→ equilibrium

~~acid~~ or ~~base~~ ⟶ 100% ions
no equilibrium...

acid or base ⟷ < 100% ions
equilibrium
ICE

<u>Note:</u>
① a strong acid or a strong base will ionize completely (100%)

② an unsaturated solution of the solute will ionize completely (100%)

Base ⇒ pOH Acid → pH

$pOH = -\log[OH^-]$ $\frac{mol}{L}$

Strontium hydroxide.
Sr^{2+} OH^-

$\underline{1}Sr(OH)_{2(s)} \xrightarrow[\text{dissolving}]{} Sr^{2+}_{(aq)} + \underline{2}OH^-_{(aq)}$

molar mass Sr, 120, +2H

$\frac{5g}{100mL}$ $\frac{\text{mass}}{\text{molar mass}}$ = mole

$\frac{5g \div 121.64g/mol}{0.100 L}$ mole ratio (1:2) $[OH^-]$

$= 0.411$ mol/L $\xrightarrow{\times 2}$ 0.822 mol/L

$pOH = -\log(OH^-)$
$= -\log 0.822$
$= +0.0851$

$pH = 14 - pOH$
$= 14 - 0.0851$
$= 13.9\cancel{149}$
$= 13.9$
↑ one sig. fig

Determine the pH of 1.7 mol/L $HBr_{(aq)}$.
(strong acid)

$H\text{-}Br + H_2O \rightarrow H_3O^+ + Br^-$ K_{eq} very large

donor H+ Acid | acceptor H+ Base | conj. acid | conj. base

1.7 mol/L ⟶ 1.7 mol/L

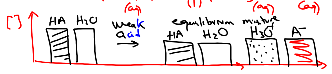
$pH = -\log(1.7)$
$= -0.23$ (2 s.f.)

Determine the pH of 1.7 mol/L CH_3COOH.
acetic acid
Weak

<u>only 6 strong acids:</u>
$HClO_4$, HBr, HCl, HI, HNO_3, H_2SO_4
because: - they ionize 100%.
- they have a very large Ka

(acid Ionization constant) $K_a = \frac{[H_3O^+][A^-]}{[HA]}$ (a ↑ acid)

$\text{acid }(HA)_{(aq)} + H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + A^-_{(aq)}$



K_a for $CH_3COOH = 1.8 \times 10^{-5}$ at 25°C
p747

<u>step 1:</u> Write the ionization equation.
Acid + Water ⟶

$CH_3COOH + H_2O \rightarrow CH_3COO^- + H_3O^+$

<u>step 2:</u> Ka of acid ⇒ 1.8×10^{-5}

<u>step 3:</u> ICE chart

	Conc.			$[H_3O^+]$
I	1.7 mol/L		0	0
C	$-x$		$+x$	$+x$
E	$1.7 - x$		x	x

<u>Step 4:</u> Solve for x

$\frac{1.7}{1.8 \times 10^{-5}} \gg 1000$

Consider the approximation method.
$[acid] / K_a > 1000$

$$\frac{x^2}{(1.7 - \cancel{x})} = K_a = 1.8 \times 10^{-5}$$
↳ approximation

$x = \sqrt{1.7(1.8 \times 10^{-5})}$
$x = 0.0055317$

<u>Step 5:</u> $[H_3O^+]_{equilib.} = x = 0.0055317$
$pH = -\log(H_3O^+)$
$= 2.2517$

How many s.f. do I need to keep? 2
$pH = 2.25$

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Apr 5-1:16 PM